From Le Châtelier to homeostasis. Osmolarity, pH in biochemistry. Buffer systems.

From the Chemistry Exam to the Final Exam in Biochemistry

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Colligative properties

A material's property is called *colligative*, if it depends only on the number of particles, but not on the type (or chemical composition) of the particles.

Comparing the properties of a pure solvent with those of a solution:

- Vapor pressure of solution is lower
- Freezing point of solution is lower
- Boiling point of solution is higher
- Osmosis

Osmosis - the physical phenomenon

- **Osmosis** (in general): Solvent movement from the high concentration site to the low concentration site through a semipermeable membrane
- The **osmotic pressure** of a solution: the pressure that needs to be expended to prevent the ingress of water

Osmotic pressure

$$\Pi = i\ c\ R\ T$$

- Π = osmotic pressure
- i = van't Hoff factor
- c = molarity of **all dissolved** substances, so $i \cdot c = i_1 \cdot c_1 + i_2 \cdot c_2 + i_3 \cdot c_3 + \dots i_n \cdot c_n$
- R = universal gas constant
- T = temperature

Question

What's the difference between molarity and osmolarity (osmotic concentration?

- Molarity (c): mol of solute in 1 dm^3 solution
- Osmolarity (c_{osm}): molarity of **all dissolved** substances, so

$$i \cdot c = i1 \cdot c1 + i2 \cdot c2 + i3 \cdot c3 + \ldots in \cdot cn$$

The difference between molarity and osmotic concentration can be illustrated by this example:

- Molarity: a 100 millimolar sodium chloride solution contains 0.1 mol NaCl per liter (c = 0.1 mol / l = 100 mmol / l).
- Osmotic concentration: in the same solution sodium chloride dissociates into the ions Na^+ and Cl^-, so 0.2 mol of osmotically active particles are dissolved (c_{osm} = 0.2 osmol / l = 200 mosmol / l). The actual osmotic concentration is slightly lower because not all particles dissociate and the solubility is temperature dependent.

Question

What is the normal osmotic concentration of blood plasma?

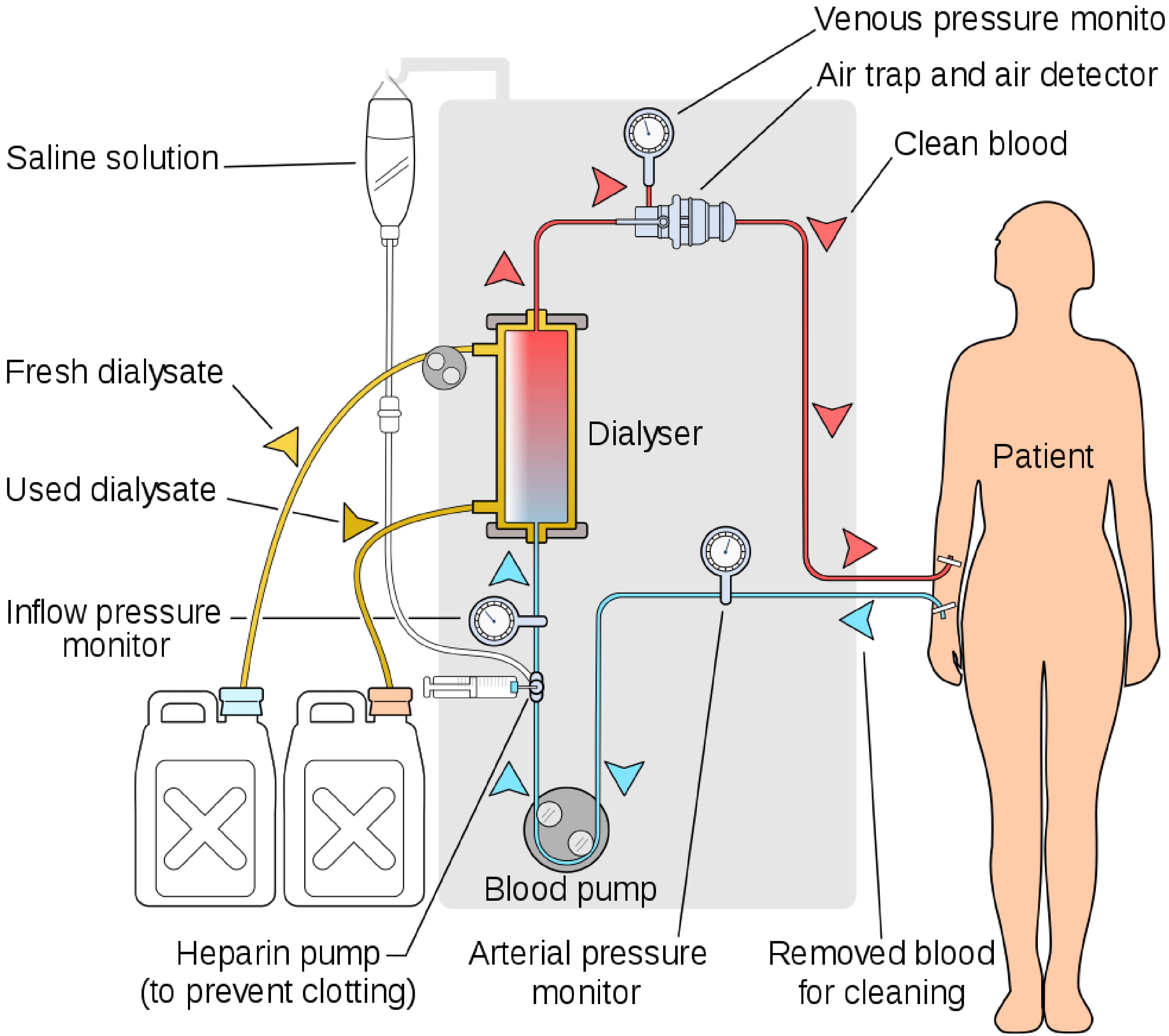
Osmolarity of blood plasma:

290 milliosmol <cBP <310 milliosmol equivalent to ~ 150 mM NaCl (0.9%)

(also known as isotonic or physiological saline)

Milliosmol = mOsm = mosmol/l

Hemodialysis



Question

Compared to water, the environment on the inside of the blood cell could best be described as:

a. hypertonic

b. ginandtonic

c. isotonic

d. hypotonic

Question

Solution	Hypotonic	Isotonic	Hypertonic
0,3M Sodium chloride			
0,05M Glucose			
0,1M Calcium chloride			
0,1M Sodium phosphate			
0,15M Urea			
0,15M Potassium chloride			
0,1M Mannose			
0,1M Magnesium chloride			
0,3M Lactate			
0,1M Ethanol			

Question

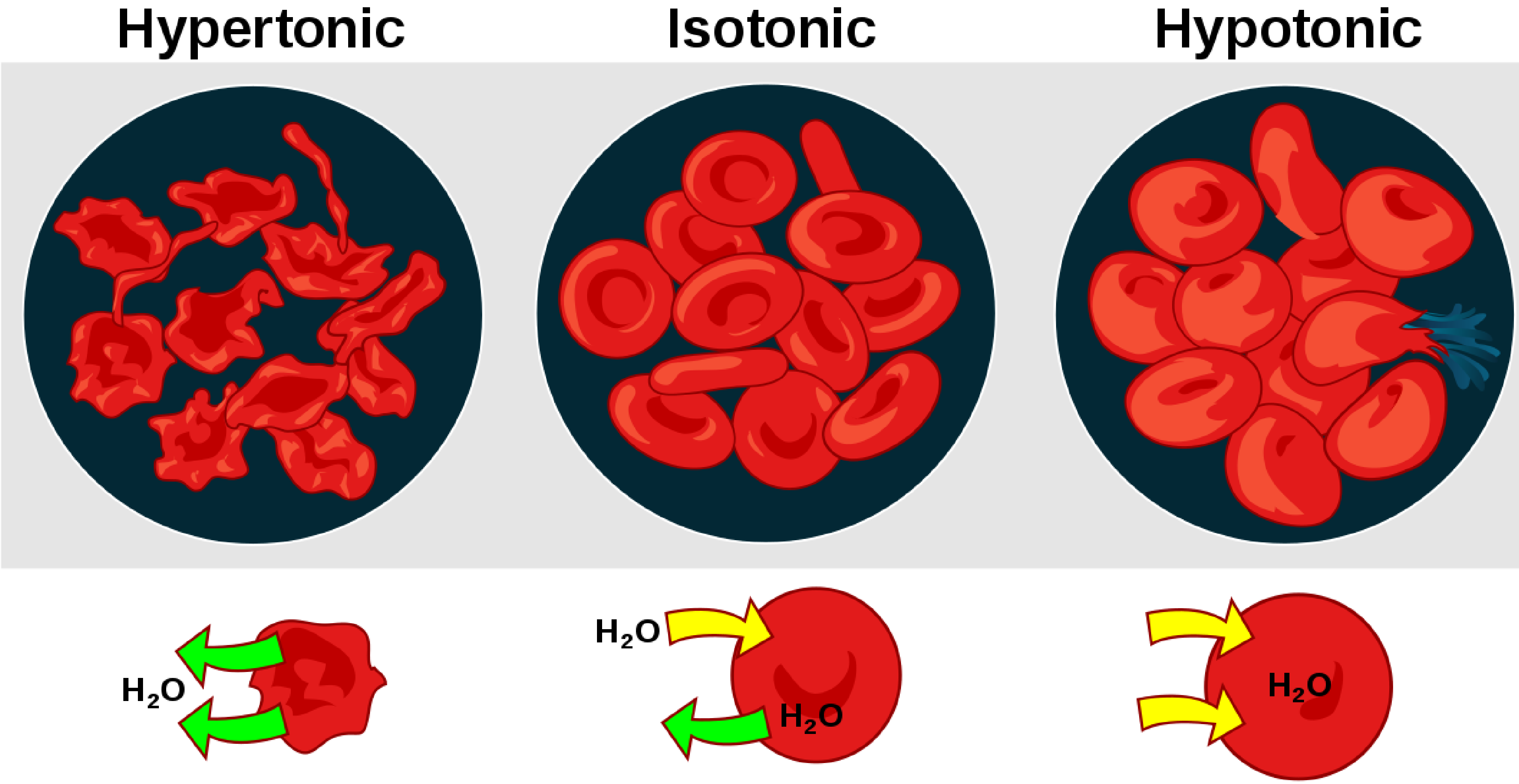
Isotonic saline solutions are preferred over hypertonic or hypotonic saline solutions when treating dehydration because __________.

A. Isotonic saline solutions do not cause cells to shrink from further dehydration, nor do they cause them to swell from over-hydration.

B. Isotonic saline solutions cause cells to shrink, making it more difficult to over-hydrate them.

C. Isotonic saline solutions cause cells to swell, making it more difficult to dehydrate them again.

D. Isotonic saline solutions are not preferred over any other saline solution.


Hypertonic
Isotonic
Hypotonic
H2O
H2O
H2O
H2O

Question

What would happen if you gave a patient an IV of pure water?

a. Their blood cells would shrink.

b. Their blood cells would burst.

c. The patient would slowly become rehydrated.

d. I would be promoted for my outstanding level of medical care.

Question

In the scenario described in the previous question, which way would water be moving?

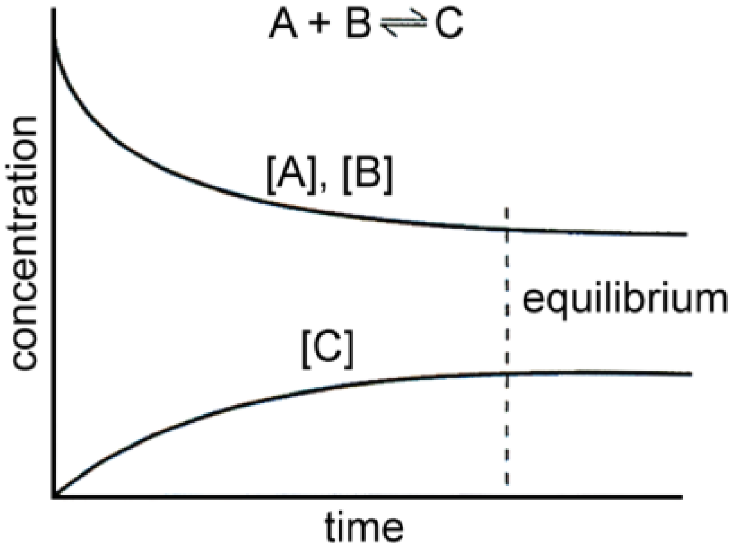
a. into the blood cells

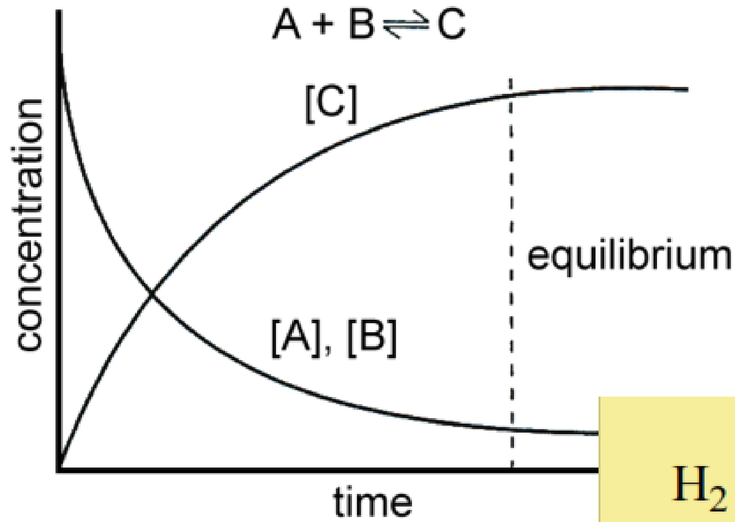
b. out of the blood cells

c. both into and out of, but with no net change

d. water would not be moving

Chemical equilibria





Equilibrium is on the side of the reactants or on the side of the product.
Reversible, dynamic, the reaction rates of the forward and the reverse reaction are the same.
Mass action law gives the value of the equilibrium constant (K).

$$H_2 + I_2 \underset{V_2}{\overset{V_1}{\rightleftharpoons}} 2\,HI$$

$V_1 = k_1 \cdot [H_2]\,[I_2]$

$V_2 = k_2 \cdot [HI]^2$

At equilibrium : $V_1 = V_2$

$k_1 \cdot [H_2]\,[I_2] = k_2 \cdot [HI]^2$

$$K = \frac{k_1}{k_2} = \frac{[HI]^2}{[H_2][I_2]}$$

The Equilibrium Law

Le-Châtelier's principle

If a system in equilibrium experiences a stress, a net reaction proceeds in the direction that relieves the stress.

Factors affecting the equilibrium:

- Change in concentration
- Change in pressure (only gas reactions)
- Change in temperature

Question

$$N_2 (g) + 3 H_2 (g) \rightleftharpoons 2 NH_3 (g)$$

- What's the type of this equilibrium?
- What happens with the equilibrium if
  - a. we increase the pressure?
  - b. we remove NH_3?
  - c. we remove N_2?
  - d. we add catalyst?

$$2Hg (l) + O_2 (g) \rightleftharpoons 2 HgO (s)$$

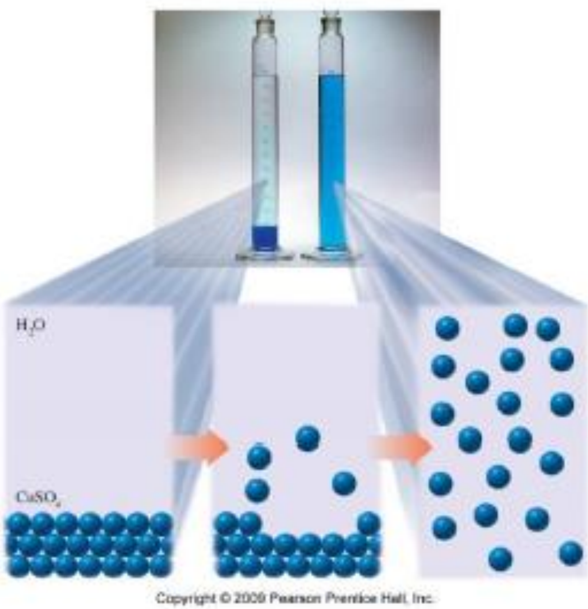
$$\Delta H° = -180 \text{ kJ}$$

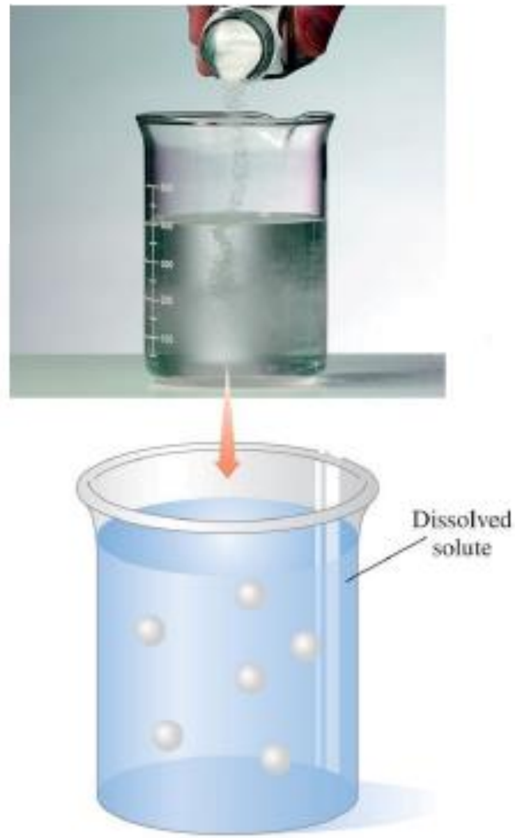
$[O_2]\uparrow$ When is this right?

Solubility

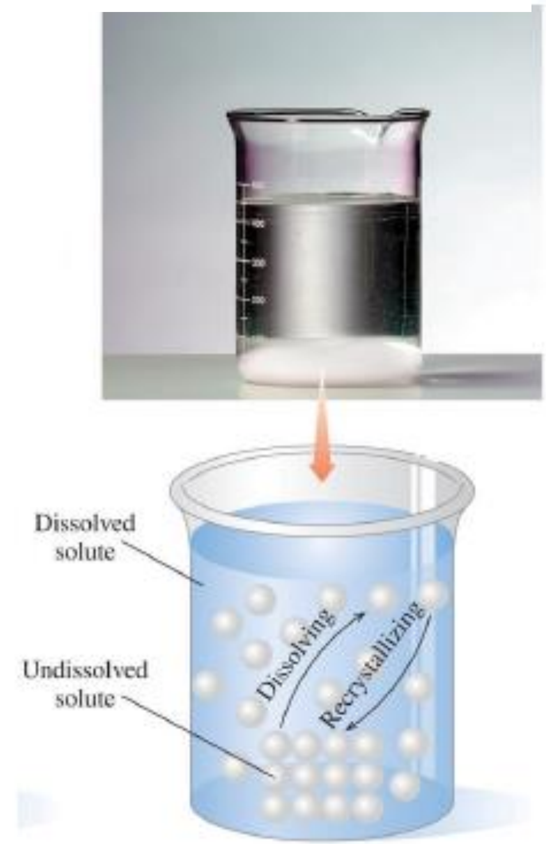
The maximum amount of a solute that can be dissolved in a certain amount of solvent at a given temperature. The concentration of the saturated solution at a given temperature.

Depends on temperature and on pressure (by gases).





Not saturated solution



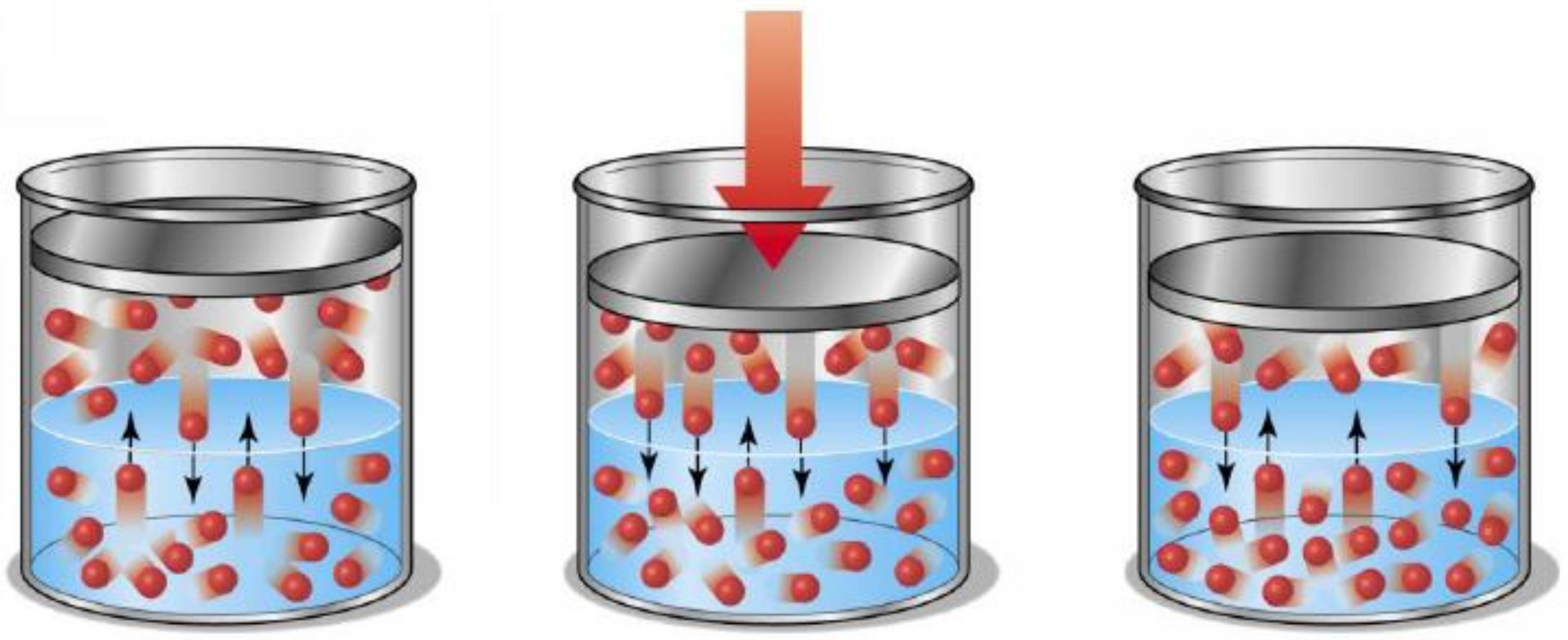
Saturated solution

Effect of pressure on the solubility of gases

Henry's Law

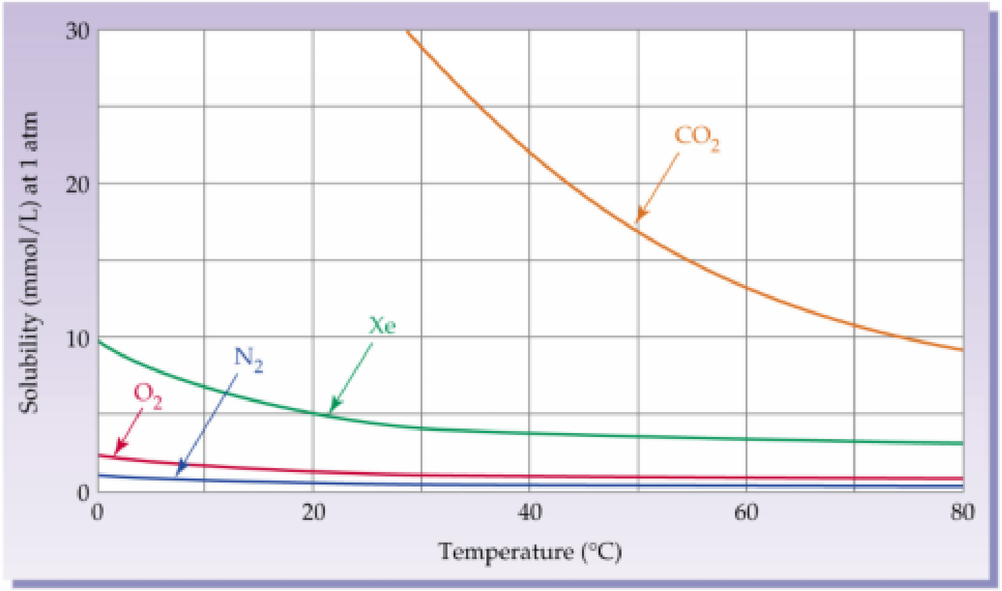
Solubility $= k \cdot p$

p: partial pressure of the gas over the solution
k: constant,
depending on quality and on temperature



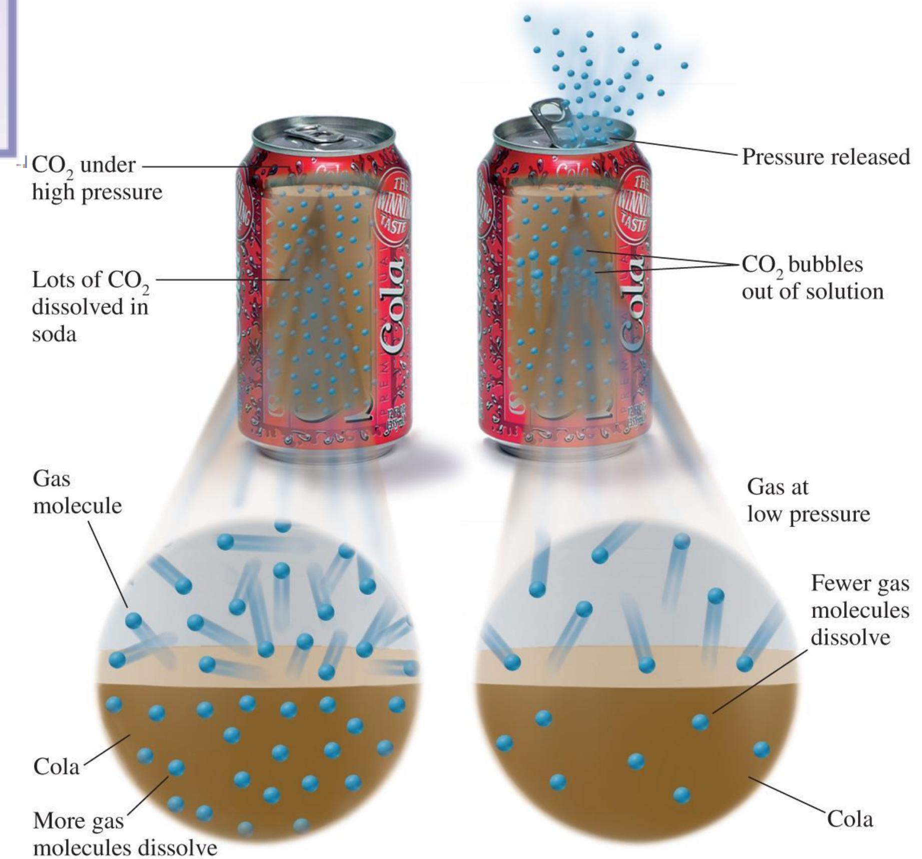
Equilibrium　　Pressure increased　　New equilibrium

Solubility of gases: inversely proportional with the temperature



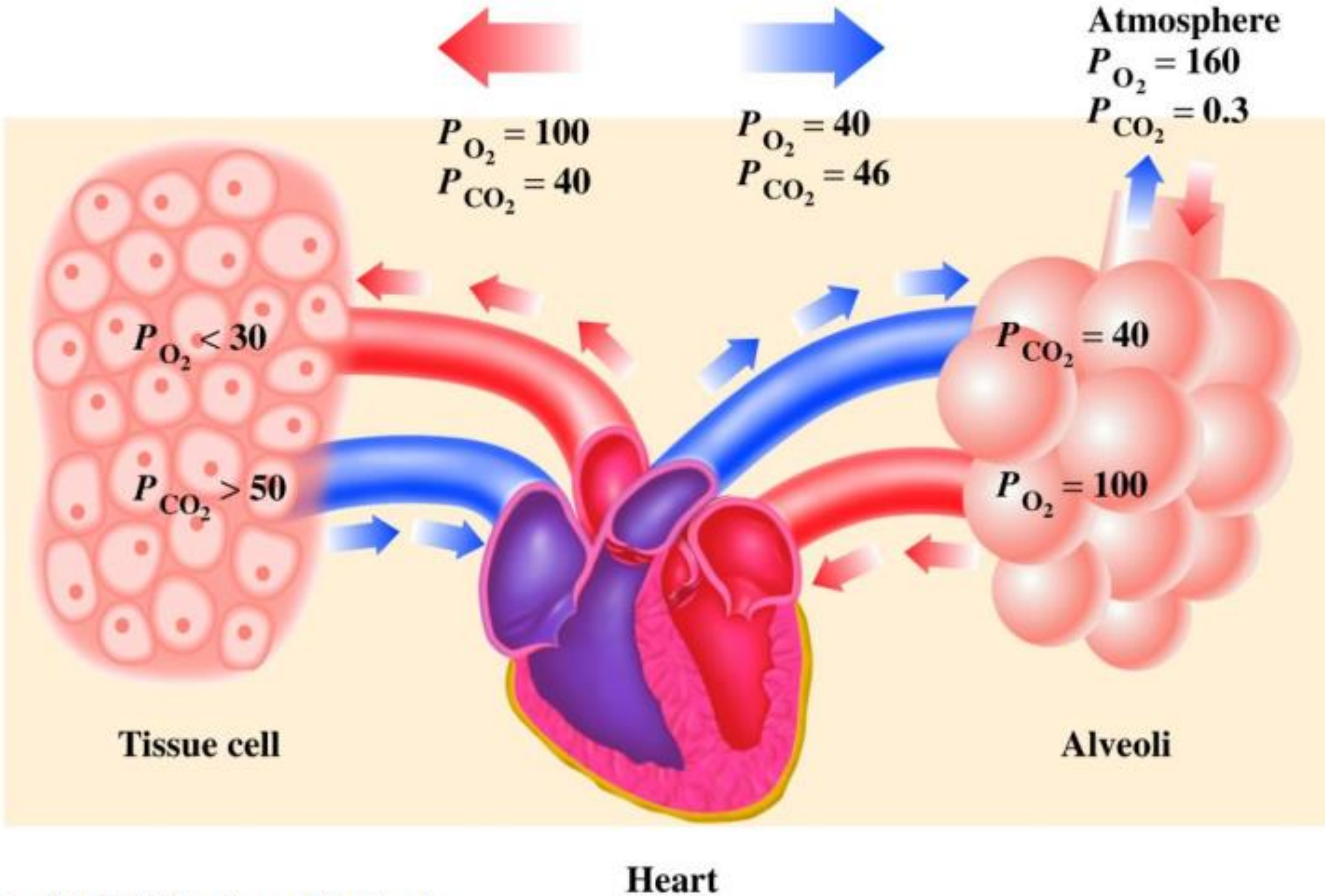
Question

Why do we drink coke, beer, champagne, etc. cold, but tea hot?



- In the lungs, O_2 enters the blood, while CO_2 from the blood is released.
- In the tissues, O_2 enters the cells, which release CO_2 into the blood.

Blood Gases



Electrolytes

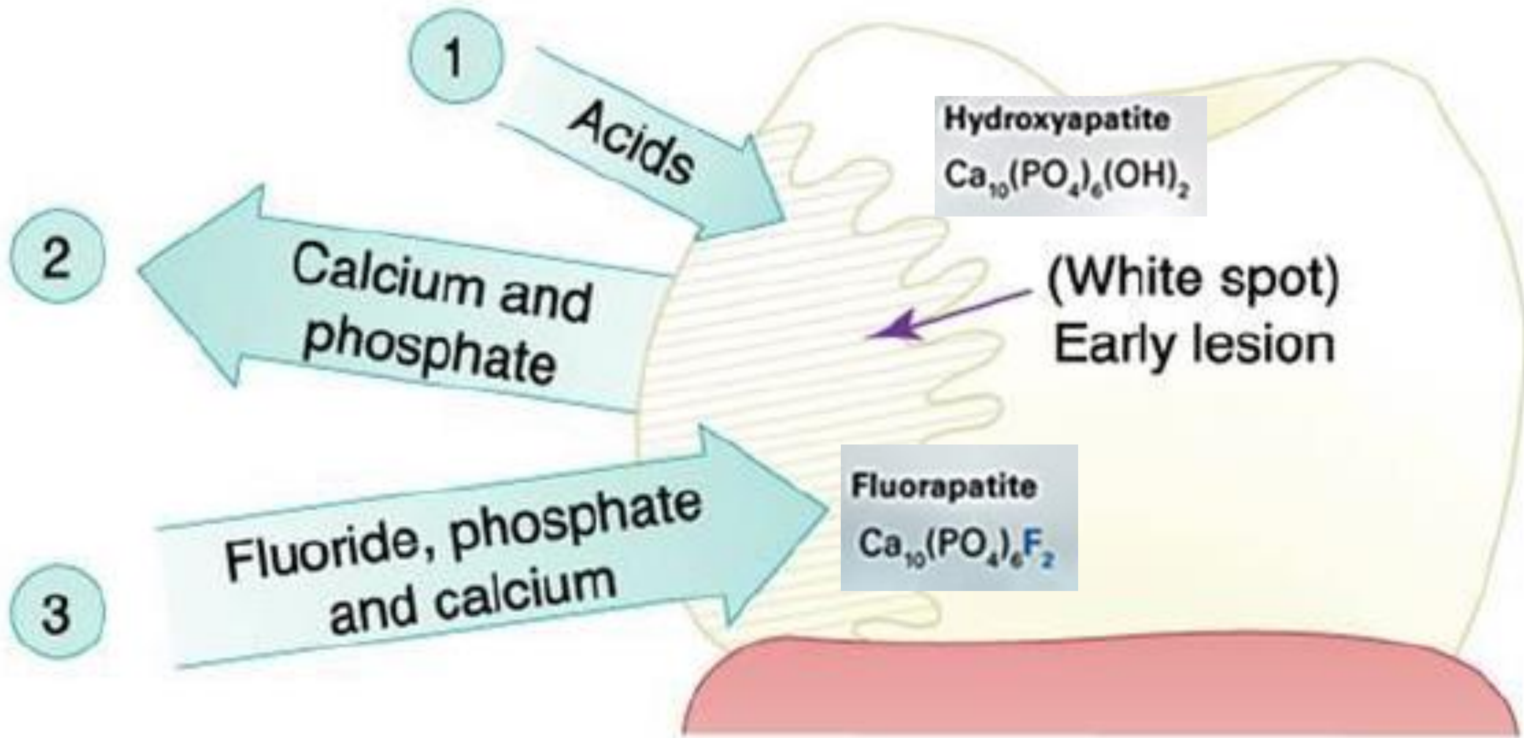
Electrolytes: ionic or polar covalent compounds.
Electrolytic dissociation: release of ions of an electrolyte when dissolved in water. (From the ion lattice, ion-dipole attractions, hydration shell)

Strong electrolytes: dissociate in large extent (70-100%), that's why they have electrical conductivity. e.g. strong acids, strong bases, water-soluble salts
In solid state there are no ions that could move freely.

Weak electrolytes: dissociate in slight extent, only partially, so they predominantly exist as molecules in solutions. e.g. weak acids, weak bases

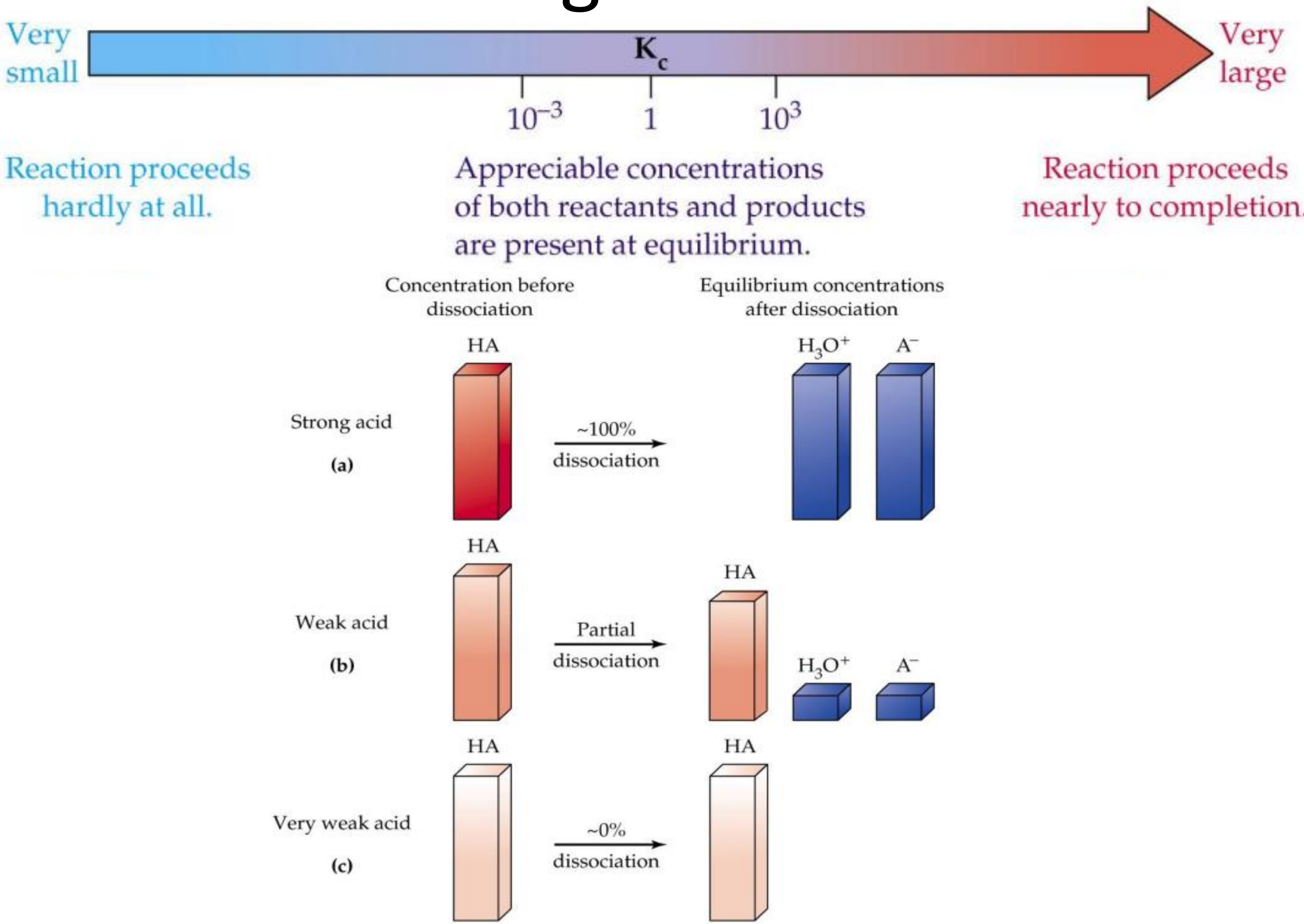
Non-electrolytes: in solutions do not form ions, therefore they do not conduct electricity at all. e.g. ethanol, sucrose, urea

Solubility: the concentration of the saturated solution means the solubility of the given substance.
It depends on the **temperature**.
Ionic compounds with basic anions have better solubility at a **lower pH**.



1. The tooth is attacked by acids in plaque and saliva.
2. Calcium and phosphate dissolve from the enamel in the process of demineralization.
3. Fluoride, phosphate and calcium re-enter the enamel in a process called remineralization.

Strong or weak?



Weak electrolytes

While they do not dissociate completely in the solution, the dissolved molecule is in equilibrium with the ions.

Degree of dissociation (α)

A fraction of the total amount of the electrolyte that is dissociated into ions. (Sometimes given in percentage.) It always depends on the concentration.

$\alpha =$ concentration of dissociated moles/concentration of total moles

If **HA** is a weak acid, and the initial concentration of the acid is ***c*** (mol/dm^3), the dissociation constant (K_a) is equal to:

$$\frac{\alpha^2 \cdot c}{1-\alpha}$$

where 1-α is approx. 1

Law of mass action

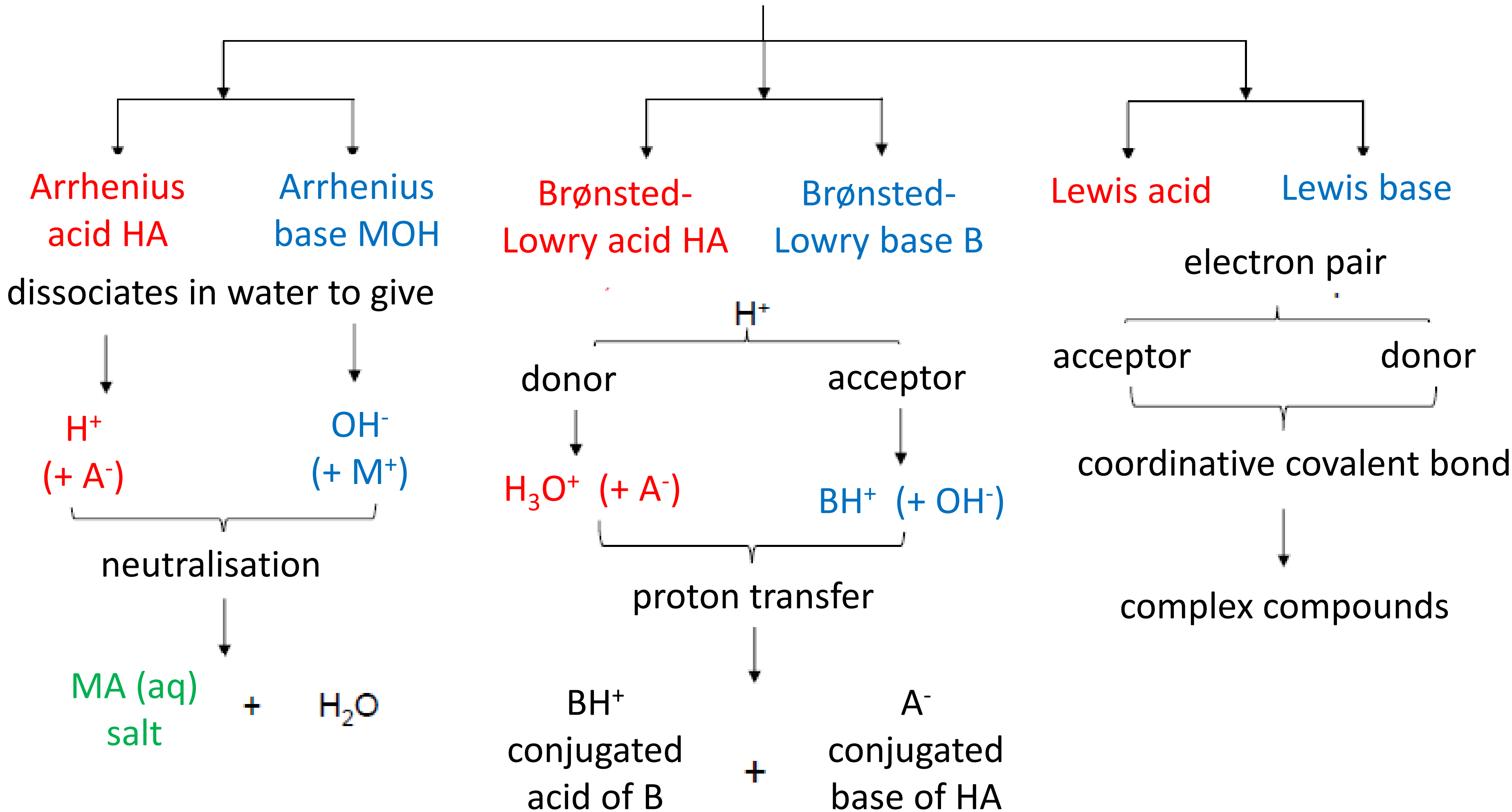
$$HCOOH \rightleftharpoons HCOO^- + H^+$$

$$K_a = \frac{[HCOO^-][H^+]}{[HCOOH]}$$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

ACIDS AND BASES



They can be:

- Strong or weak
- Monoprotic or polyprotic

Arrhenius theory:

Acids dissociate in water to produce H^+ or H_3O^+ ions

$$HA\ (aq) \rightleftharpoons H^+\ (aq) + A^-\ (aq)$$ e.g. HCl, HCN, HNO_3, CH_3-COOH

Bases dissociate in water to generate OH^- ions

$$MOH\ (aq) \rightleftharpoons M^+\ (aq) + OH^-\ (aq)$$ e.g. NaOH, NH_3

Brønsted-Lowry theory:

Acids are proton donor molecules or ions e.g. NH_4^+, HSO_4^-

Bases are proton acceptor substances e.g. CN^-

Acid-base reactions are proton transfer reactions.

An acid losing a proton produces a *conjugated base*, a base accepting a proton yields a *conjugated acid*.

Lewis theory: (dative or coordinative bond)

Lewis acids: electron-pair acceptors

Lewis bases: electron-pair donors

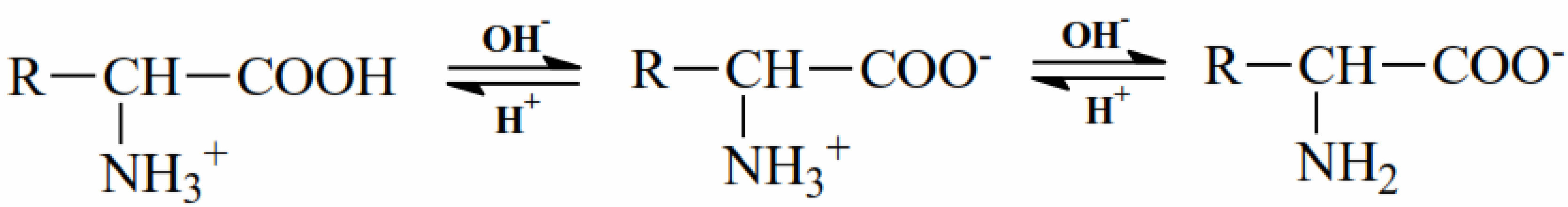
Amphoterism

Definition: they can behave both as an acid and as a base depending on the partner.

Inorganic amphoteric compounds

Amphoteric oxides: oxides not soluble in water, but soluble in acids an bases (e.g. Al_2O_3, ZnO).

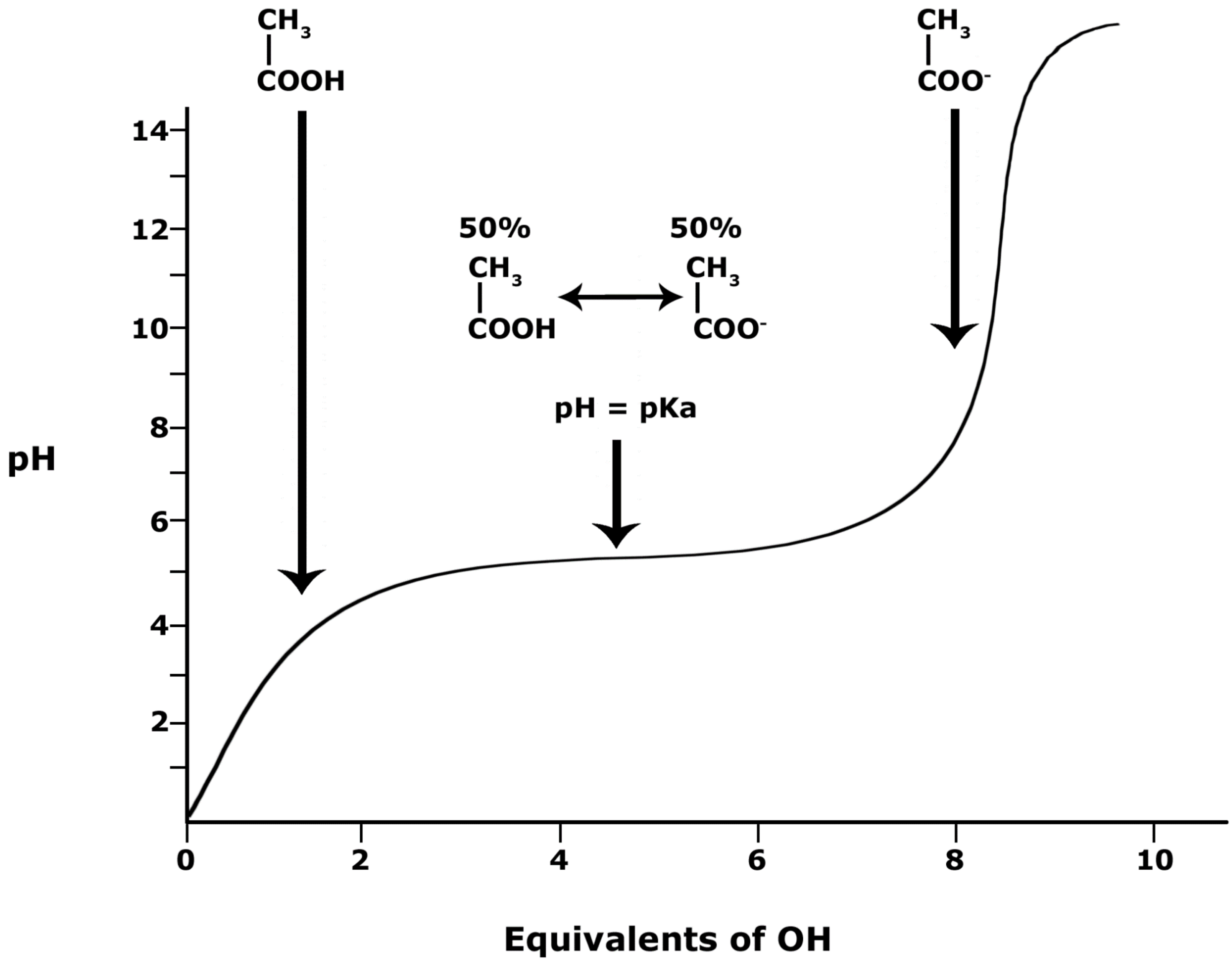
Amphoteric character of amino acids



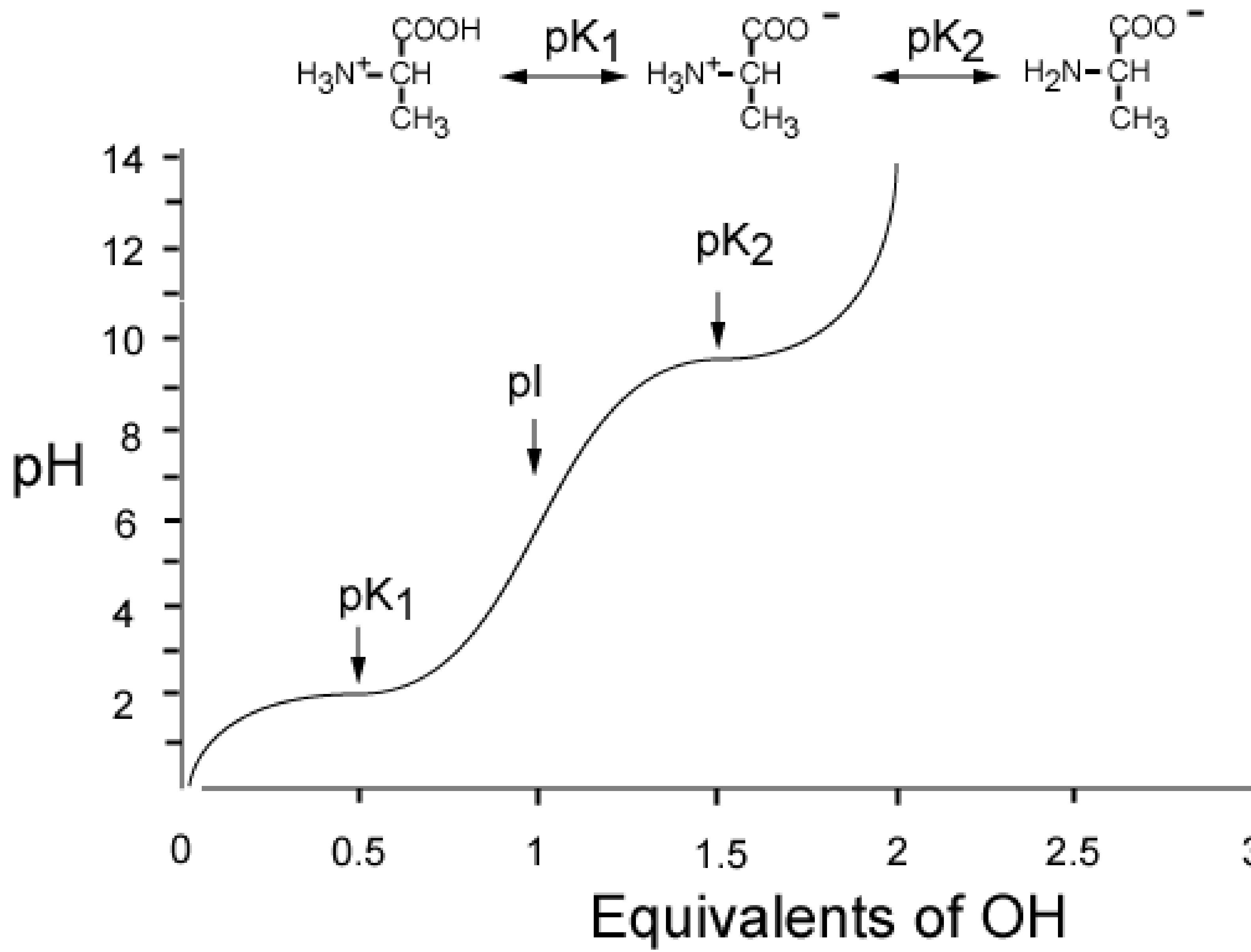
cation **(acidic medium)**

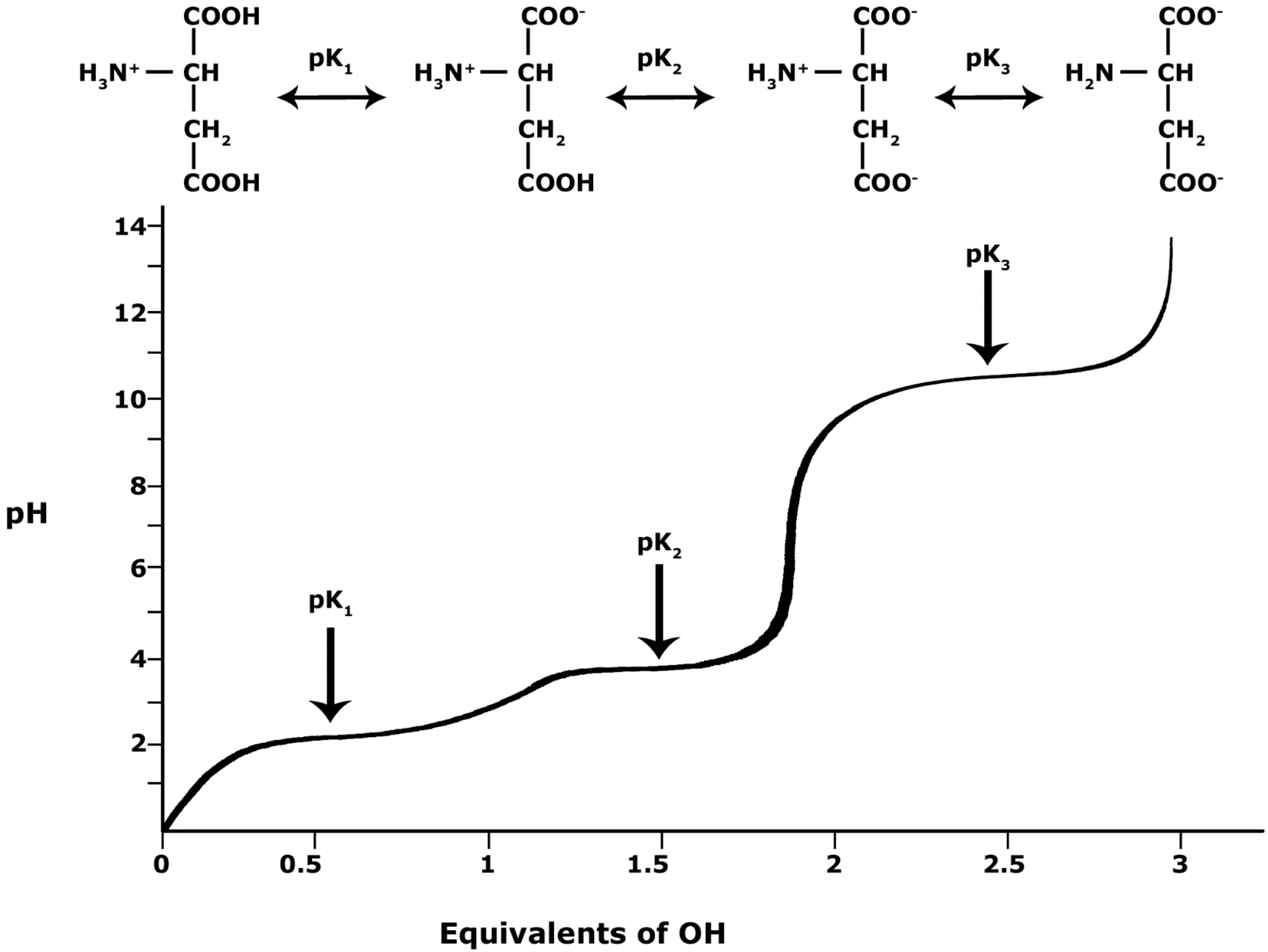
zwitterion **(~ neutral medium)**

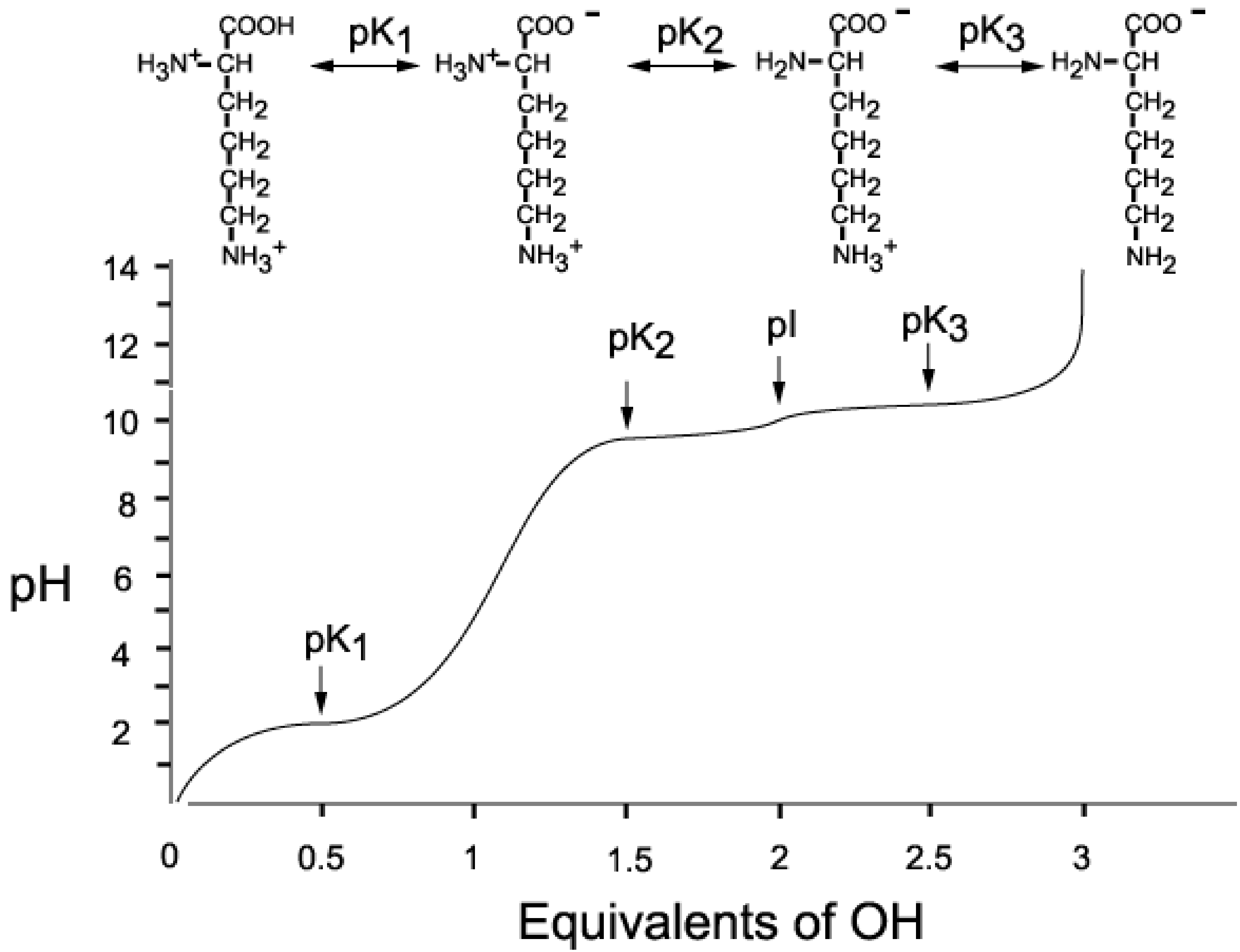
anion **(basic medium)**



CH3
COOH
CH3
COO-
50%
CH3
COOH
50%
CH3
COO-
pH = pKa
pH
14
12
10
8
6
4
2
0
2
4
6
8
10
Equivalents of OH


COOH
H3N+- CH
CH3
pK1
COO -
H3N+- CH
CH3
pK2
COO -
H2N- CH
CH3
14
12
10
8
6
4
2
pH
pK2
pI
pK1
0
0.5
1
1.5
2
2.5
3
Equivalents of OH

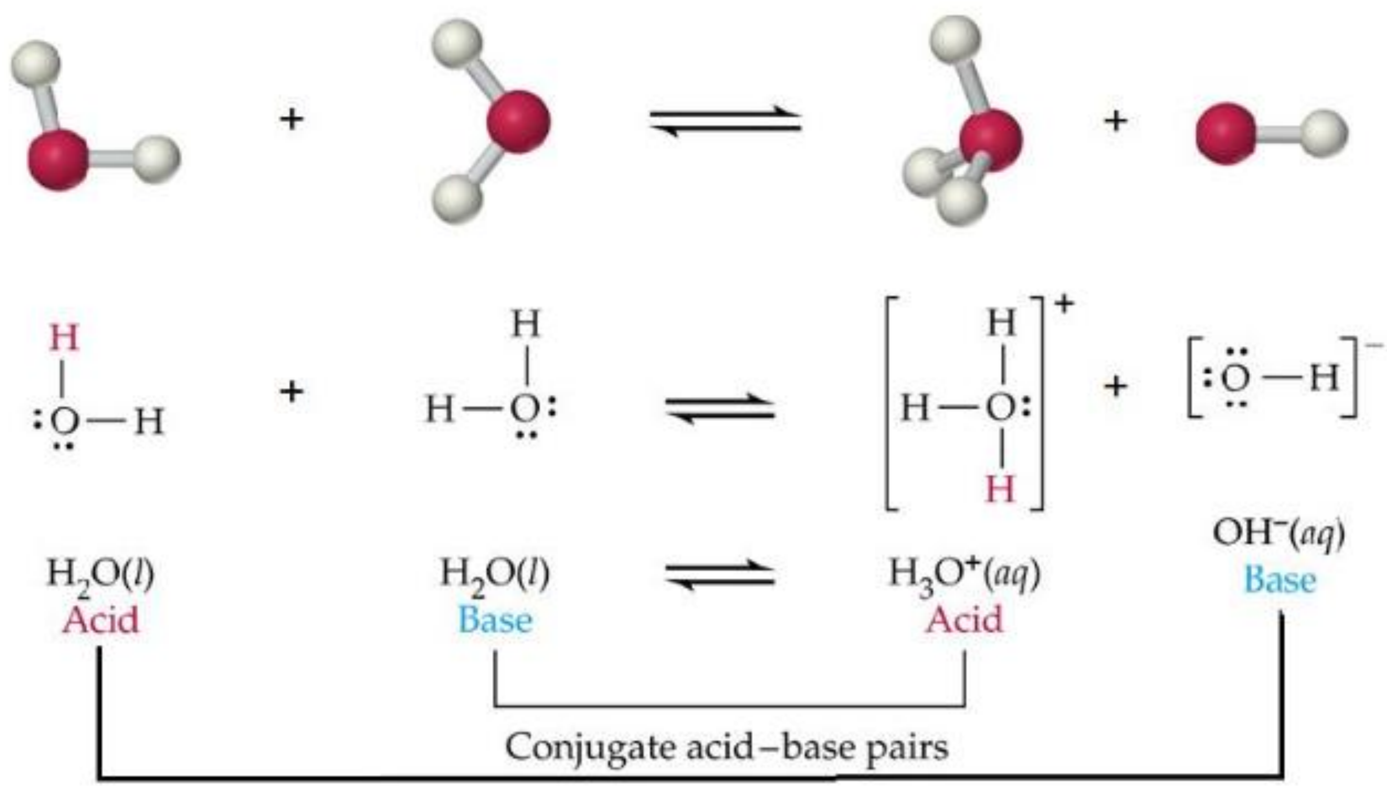

COOH
H3N+ — CH
CH2
COOH
pK1
COO-
H3N+ — CH
CH2
COOH
pK2
COO-
H3N+ — CH
CH2
COO-
pK3
COO-
H2N — CH
CH2
COO-
14
12
10
8
6
4
2
pH
pK1
pK2
pK3
0
0.5
1
1.5
2
2.5
3
Equivalents of OH



COOH
H_3N^+-CH
CH_2
CH_2
CH_2
CH_2
NH_3^+
pK_1
COO^-
H_3N^+-CH
CH_2
CH_2
CH_2
CH_2
NH_3^+
pK_2
COO^-
H_2N-CH
CH_2
CH_2
CH_2
CH_2
NH_3^+
pK_3
COO^-
H_2N-CH
CH_2
CH_2
CH_2
CH_2
NH_2
14
12
10
8
6
4
2
pH
pK_1
pK_2
pI
pK_3
0
0.5
1
1.5
2
2.5
3
Equivalents of OH

The self-ionization of water

The water behaves as a very weak electrolyte showing small conduction. It is an **amphoteric compound** having **both acidic and basic properties**.



$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+ + OH^-$$

Dissociation of water

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+ + OH^-$$

Applying the law of mass action:

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2} = 1,8 \times 10^{-16}$$

Because the water is a very weak electrolyte, the concentration of the water remains constant, it is incorporated into the equilibrium constant.

Ion-product constant of water

$$K_W = [H^+]\,[OH^-] = 1.00 \cdot 10^{-14} \quad (\text{at } 25\ ^0\text{C})$$

The temperature influences the dissociation.

In pure water: $[H^+] = [OH^-] = 1.00 \cdot 10^{-7}\ mol/dm^3$

The pH and the pOH

pH is the negative logarithm of the hydronium ion concentration.

$$pH = -\log [H_3O^+]$$

$$pH = -\log [H^+] = \log \frac{1}{[H^+]}$$

pOH is the negative logarithm of the hydroxide ion concentration

$$pOH = -\log [OH^-] = \log \frac{1}{[OH^-]}$$

The connection between the pH and pOH is the following:

$$K_w = [H^+]\,[OH^-] = 1.00 \cdot 10^{-14} \quad (\text{at } 25\ ^0C)$$

$$pK_w = -\log K_w = 14.00$$

$$pK_w = \mathbf{pH + pOH = 14.00}$$

In neutral solution:

$$\mathbf{pH = pOH = 7.00}$$

Hydrolysis of salts

Definition: the reaction of ions with water to yield protons or hydroxide ions.

- Salts of strong acids and strong bases have always a neutral solution. (pH=7). E.g. NaCl, KCl
- Salts derived from a strong base and a weak acid (e.g. CH_3-COONa) mean anion hydrolysis in the solution (pH>7). The weak acid is protonated, so the OH^- concentration increases and the solution is basic.
- Salts derived from a weak base and a strong acid (e.g.NH_4Cl) mean cation hydrolysis in the solution (pH<7). The weak base is deprotonated, the H^+ concentration increases, the solution is acidic.
- Salts derived from a weak base and a weak acid (e.g. NH_4NO_2) mean both anion and cation hydrolysis simultaneously in the solution. The pH value is determined by the relative magnitude of the K values. (In this case K_a of NH_4^+ ions is larger than K_b of NO_2^- ions, so the solution is acidic.)

Buffer solutions

Definition: Buffer solutions contain a mixture of a weak acid and its conjugate base (or salt) or a weak base and its conjugate acid (or salt).

They moderate the effects of acids or bases on pH change: the salt causes higher concentration of the conjugate base/acid than the weak acid/base alone. The buffer solutions contain approximately the same concentration of the weak acid/base and its salt.

e.g. CH_3-COOH + CH_3-COONa + acid $\longrightarrow$ salt absorbs H^+

+ base $\longrightarrow$ acid neutralizes

General expression for acidic buffers:

$$[H^+] = K_a \cdot \frac{[HA]}{[A^-]}$$

$$-\log [H^+] = -\log K_a + \log \frac{[A^-]}{[HA]}$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

The equation above is the **Henderson-Hasselbalch equation** describing the **pH value of an acidic buffer system, which depends only on the ratio of the components**.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$[OH^-] = K_b \frac{[NH_3]}{[NH_4^+]} \qquad pOH = pK_b + \log \frac{[NH_4^+]}{[NH_3]}$$

$$pOH = pK_b + \log \frac{[HB^+]}{[B]}$$

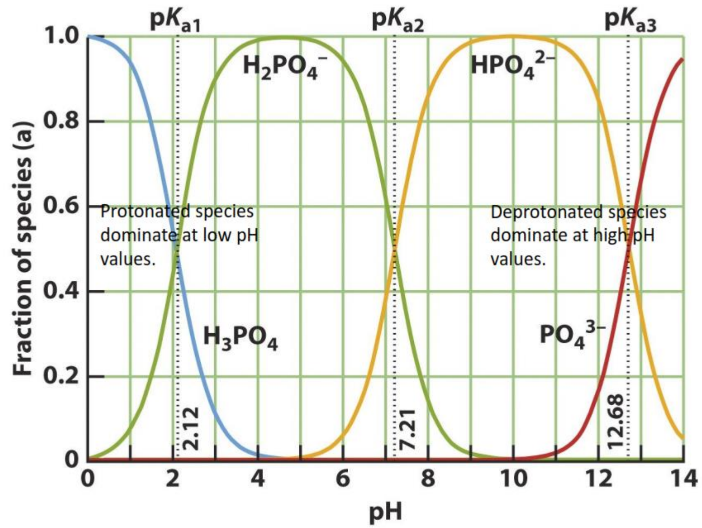
HB^+ = conjugated acid

B = base

Henderson-Hasselbalch equation for basic buffers.

The most important buffer systems in humans:

- Phosphate buffer in the intracellular space
- Carbonic acid / bicarbonate system in the blood and in interstitial space
- Albumin in blood plasma
- Hemoglobin in red blood cells



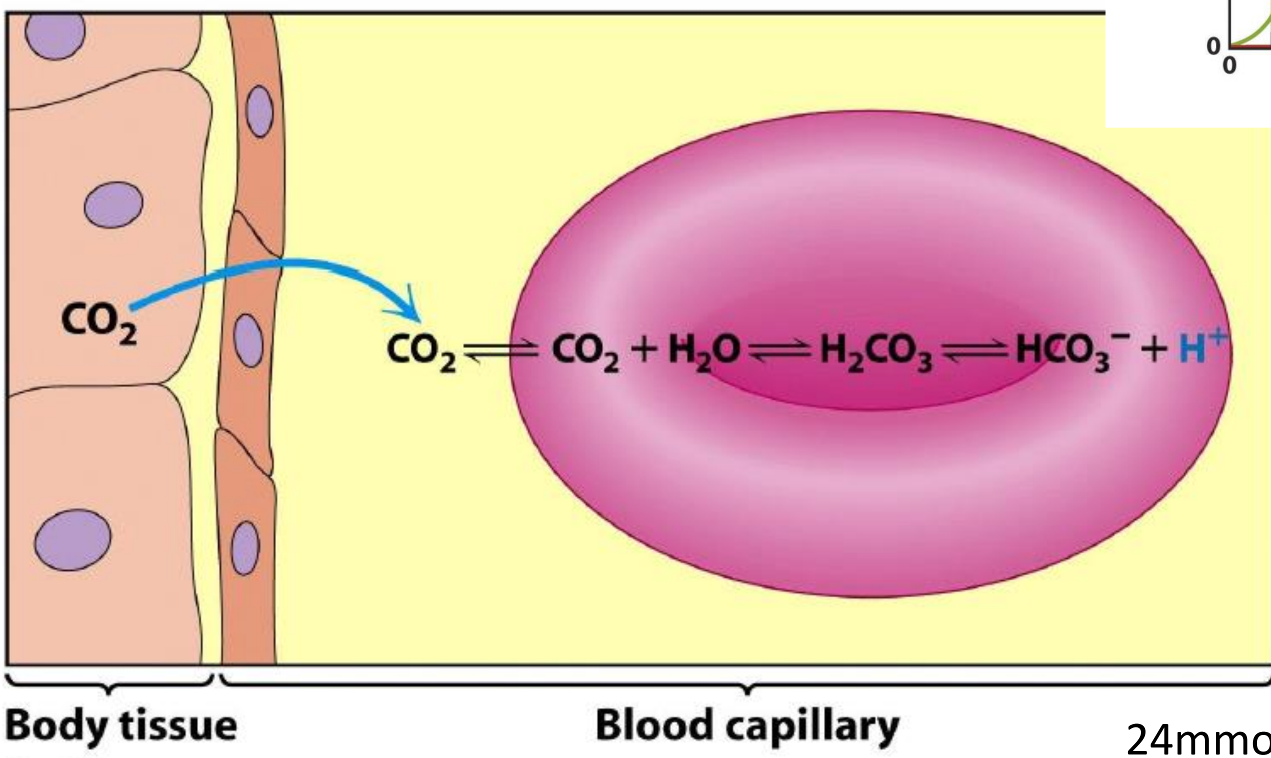


Figure 7-20
Biochemistry, Sixth Edition

24mmol/dm^3

$$pH = pKa + \log \frac{[HCO_3^-]}{[H_2CO_3]} = 6{,}1 + \log \frac{[HCO_3^-]}{0{,}03pCO_2} = 6{,}1 + \log 20 = 7{,}4$$

Solubility of CO_2

Buffer capacity

Definition:
The amount of strong acid or strong base in moles required to shift the pH of 1l buffer solution by one unit.

Question

Which of the following buffer solutions has a higher buffer capacity against the 0.1 M NaOH solution?

a. 0,1 M Na_2HPO_4 and 0,15 M NaH_2PO_4
b. 0,15 M Na_2HPO_4 and 0,1 M NaH_2PO_4

Question

Most of the fluids produced by the human body are ...

a. very acidic

b. in the neutral range

c. very basic

d. very alkaline

Question

The difference between a strong acid and a weak acid involves...

a. how much of it dissociates in water

b. how much it synthesizes in water

c. how much of it changes into water

d. hours spent lifting weights

Question

Each number on the pH scale represents a ________ change in H^+ concentration.

a. one point

b. tenfold

c. hundredfold

d. ten thousandfold

e. negative

Question

A salt is:

a. A compound that dissociates in water and releases hydrogen ions (H^+).

b. An atom that dissociates in water and forms cations and anions.

c. A compound that dissociates in water and releases hydroxyl ions (OH^-).

d. A compound that dissociates in water and produces cations or anions other than H^+ or OH^-.

Question

The body's bicarbonate buffer system functions to keep the pH of the blood neutral. Which two organs/systems are primarily involved in keeping this buffer system balanced?

a. lungs, liver

b. digestive, circulatory

c. digestive, respiratory

d. kidneys, lungs

e. immune, circulatory

Redox reactions: chemical reactions with electron transfer

e.g.: $\overset{0}{C} + \overset{0}{O_2} \rightarrow \overset{+4}{C}\overset{-2}{O_2}$

$\underline{\overset{0}{Zn}} + 2\ H\overset{+1}{Cl} \rightarrow \overset{+2}{Zn}Cl_2 + \overset{0}{H_2}$

$2\ H_2\overset{-2}{S} + \overset{+4}{S}O_2 \rightarrow 2\ \underline{\overset{0}{S}} + 2\ H_2O$ (synproportionation)

$\overset{0}{Cl_2} + 2\ NaOH \rightarrow 2\ Na^+ + \overset{-1}{Cl^-} + O\overset{+1}{Cl^-} + H_2O$ (disproportionation)

Oxidation: loss of electrons (increase of oxidation number)

Reduction: take-up of electrons (decrease of oxidation number)

Non-redox reaction: $\underline{CaCO_3} \rightarrow \underline{CaO} + CO_2$

$NH_3 + H^+ \rightarrow NH_4^+$ (acid-base reactions)

$Al^{3+} + 4\ OH^- \rightarrow [Al(OH)_4]^-$ (complex formations)

$AgNO_3 + NaCl \rightarrow \underline{AgCl} + NaNO_3$

TABLE 12.1 Standard Potentials at 25°C*

More p**O**sitive → **O**xidizes more negative

strong oxidizing agents

weak reducing agents

Species	Reduction half-reaction	E° (V)
	trongly oxidizing	
	$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$	
	$Au^+(aq) + e^- \longrightarrow Au(s)$	
Ce^{4+}/Ce^{3+}	$Ce^{4+}(aq) + e^- \longrightarrow Ce^{3+}(aq)$	+1.61
$MnO_4^-, H^+/Mn^{2+}, H_2O$	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.51
Cl_2/Cl^-	$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}, H^+/Cr^{3+}, H_2O$	$Cr_2O_7^{2-} + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$O_2, H^+/H_2O$	$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	+1.23; +0.82 at pH
Br_2/Br^-	$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq)$	+1.09
$NO_3^-, H^+/NO, H_2O$	$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$	+0.96
Ag^+/Ag	$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.80
Fe^{3+}/Fe^{2+}	$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.77
I_2/I^-	$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.54
$O_2, H_2O/OH^-$	$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40; +0.82 at pH
Cu^{2+}/Cu	$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.34
$AgCl/Ag, Cl^-$	$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$	+0.22
H^+/H_2	$2H^+(aq) + 2e^- \longrightarrow H_2(g)$	0, by definit
Fe^{3+}/Fe	$Fe^{3+}(aq) + 3e^- \longrightarrow Fe(s)$	−0.04
$O_2, H_2O/HO_2^-, OH^-$	$O_2(g) + H_2O(l) + 2e^- \longrightarrow HO_2^-(aq) + OH^-(aq)$	−0.08
Pb^{2+}/Pb	$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	−0.13
Sn^{2+}/Sn	$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$	−0.14
Fe^{2+}/Fe	$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	−0.44
Zn^{2+}/Zn	$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	−0.76
$H_2O/H_2, OH^-$	$2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$	−0.83; −0.42 at pH
Al^{3+}/Al	$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$	−1.66
	$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$	
	$Na^+(aq) + e^- \longrightarrow Na(s)$	
	$K^+(aq) + e^- \longrightarrow K(s)$	
	$Li^+(aq) + e^- \longrightarrow Li(s)$	
Reduced form is strongly reducing		

spontaneous

non-spontaneous

weak oxidizing agents

strong reducing agents

*For a more extensive table, see Appendix 2B.

More n**e**gative → r**e**duces more positive

Redox processes in metabolism

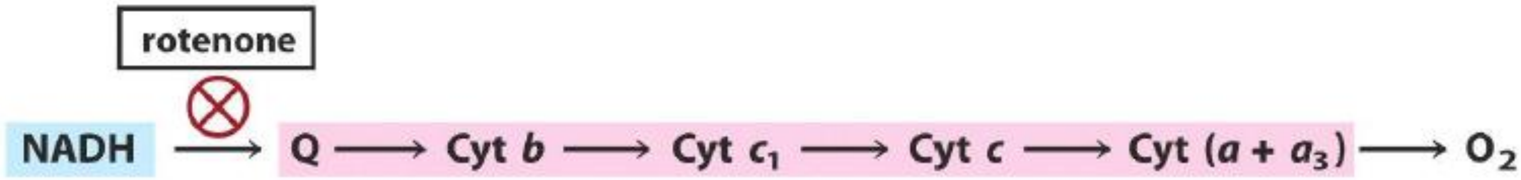
redox process (ox + *n* e^- → red)	*n*	$\varepsilon°'$ (V)
½ O_2 *(g)* + 2 H^+ + 2 e^- → H_2O *(l)*	2	0.81
cytochrome-c_1 (Fe^{3+}) → cytochrome-c_1 (Fe^{2+})	1	0.22
dehydroascorbate → ascorbate	2	0.08
fumarate → succinate	2	0.03
piruvate→ lactate	2	- 0.19
acetate → acetaldehyde	2	- 0.20
$NADP^+$ + H^+ + 2 e^- → NADPH	2	- 0.32
NAD^+ + H^+ + 2 e^- → NADH	2	- 0.32
2 H^+ + 2 e^- → H_2 *(g)* (pH = 7)	2	- 0.41
acetate → acetaldehyde	2	- 0.60
α-ketoglutarate → succinate + CO_2 *(g)*	2	- 0.67

$NAD^+ + H^+ \; 2e^- \longrightarrow NADH \qquad \varepsilon°' = -0.32\ V$

$\text{piruvate} + 2H^+ + 2e^- \longrightarrow \text{lactate} \qquad \varepsilon°' = -0.19\ V$

$$\text{piruvate} + NADH + H^+ \longrightarrow \text{lactate} + NAD^+ \qquad E°'_{el} = 0.13\ V$$

$(\Delta G°' = -nFE°'_{el} = -2 \cdot 96500 \cdot 0.13 = -25090\ J)$



redox process (ox + *n* e⁻ → red)	*n*	ε°' (V)
½ O_2 *(g)* + 2 H^+ + 2 e^- → H_2O *(l)*	2	0.81
cytochrome-a_3 (Fe^{3+}) + e^- → cytochrome-a_3 (Fe^{2+})	1	0.55
cytochrome-a (Fe^{3+}) + e^- → cytochrome-a (Fe^{2+})	1	0.29
cytochrome-c (Fe^{3+}) → cytochrome-c (Fe^{2+})	1	0.25
cytochrome-c_1 (Fe^{3+}) → cytochrome-c_1 (Fe^{2+})	1	0.22
cytochrome-b (Fe^{3+}) → cytochrome-b (Fe^{2+})	1	0.07
ubiquinone + 2H^+ + 2e^- → ubiquinol	2	0.04
NADH dehydrogenase (FMN) + 2H^+ + 2e^- → NADH dehydrogenase ($FMNH_2$)	2	-0.03
$NADP^+$ + H^+ + 2 e^- → NADPH	2	- 0.32
NAD^+ + H^+ + 2 e^- → NADH	2	- 0.32
2 H^+ + 2 e^- → H_2 *(g)* (pH = 7)	2	- 0.41

Question

Put the following components of the respiratory chain in correct order as the electrons flow in the mitochondrion!

a) Cytochrome c (Fe^{3+}/Fe^{2+}) ε^{o} = + 0.26 V
b) Q / QH_2 ε^{o} = - 0.10 V
c) Cytochrome c_1 (Fe^{3+}/Fe^{2+}) ε^{o} = + 0.23 V
d) $NAD^+/NADH$ ε^{o} = - 0.32 V
e) Cytochrome b (Fe^{3+}/Fe^{2+}) ε^{o} = + 0.04 V